1) Diesel fuel is used to heat 2.89 kg of water at 25.0 °C to a temperature of 90.0 °C. What mass, in kg, of diesel is needed to heat the water if 30.00% of the energy of combustion is allowed to escape into the environment.

Step 1 Calculate the energy needed to heat the water from 25.0 to 90.0 °C.

=> E(J) = 4.18 X 2890 X 65.0 =785 kJ

Step 2 Since this only represents 70.00% of the total energy released by the diesel calculate the total energy that must be released by burning diesel.

Allow "x" to be the total energy released.

=> 785kJ = x X 0.700

=> 785kJ / 0.700 = x = 1122kJ

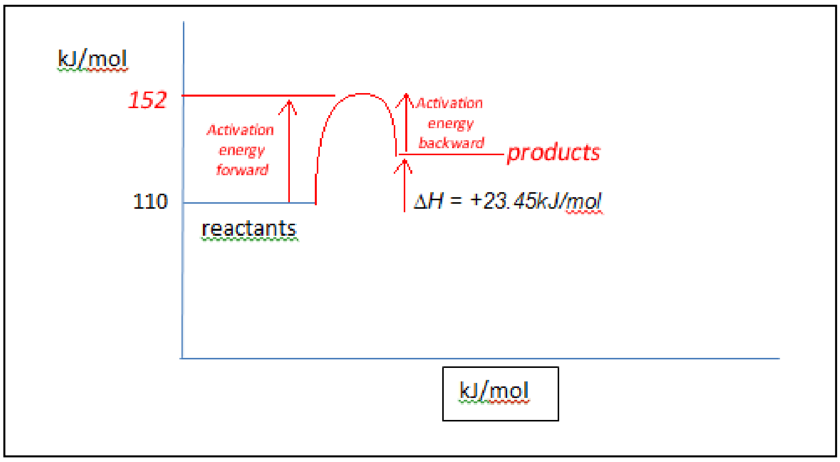
Step 3 Find how many grams of diesel will deliver 1122kJ of energy. Refer to the Data Boklet

=> 1122 kJ / 45.0 kJ/g = 24.9 grams or 0.0249 kg

2) Consider the chemical reaction given below.

$A_2(g) + B_2(g) \rightarrow 2AB(s)$ $\Delta H = +23.45kJ/mol$

The amount of energy needed to break bonds during this reaction is 42.00 kJ /mol.



i. Draw the energy profile for this reaction on the set of axes above.

Label the:

- ΔH

- Activation energy for both the forward and backward reactions. Give the magnitude of the activation energies.

  Activation energy for the forward reaction is 42.00 kJ/mol

- *Activation energy for the reverse reaction is 42.00 -23.45 = 18.55 kJ/mol*

ii. What is the energy content of the products? *133.45 kJ/mol*

iii. Consider the two chemical equations below

a) $2AB(s) \rightarrow A_2(g) + B_2(g)$ $\Delta H = ?kJ/mol$

b) $A_2(g) + B_2(g) \rightarrow 2AB(g)$ $\Delta H = ?kJ/mol$

How does the ΔH *of each of the above two reactions differ from +23.45kJ/mol. Explain*

ΔH for equation a) is -23.45 kJ/mol

ΔH for equation b) is > 23.45 kJ/mol as the product is a gas and requires more energy to keep it in the gaseous phase than its previous solid phase.